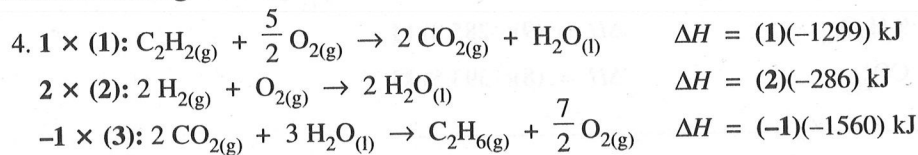


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Understanding Concepts



Since the reaction is written for one mole of ethyne, the molar enthalpy of combustion of ethyne, ΔH_{comb} , is -311 kJ/mol .

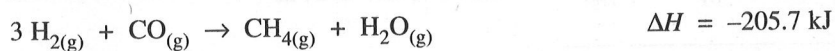
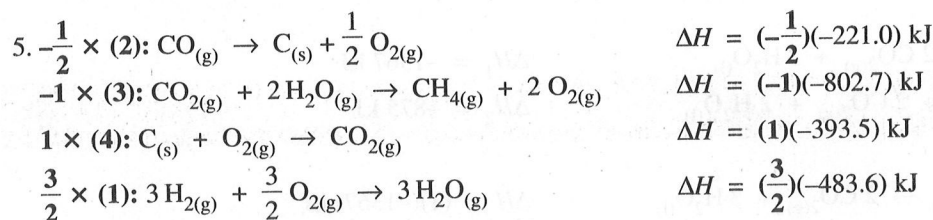
$$\text{amount of ethyne, } n = 200 \text{ g ethyne} \times \frac{1 \text{ mol}}{26.0 \text{ g}}$$

$$n = 7.69 \text{ mol}$$

$$\Delta H = n\Delta H_{\text{comb}}$$

$$= 7.69 \text{ mol} \times (-311 \text{ kJ/mol})$$

$$\Delta H = -2.39 \times 10^3 \text{ kJ}$$



Since the reaction is written for one mole of carbon monoxide, the molar enthalpy of combustion, ΔH_{comb} , is -205.7 kJ/mol CO .

$$\text{amount of CO, } n = 300 \text{ g} \times \frac{1 \text{ mol}}{28.0 \text{ g}}$$

$$n = 10.7 \text{ mol}$$

$$\Delta H = n\Delta H_{\text{comb}}$$

$$= 10.7 \text{ mol} \times (-205.7 \text{ kJ/mol})$$

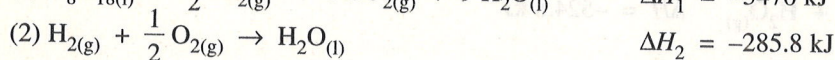
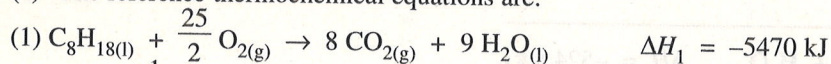
$$\Delta H = -2.20 \times 10^3 \text{ kJ}$$

SECTION 5.4 QUESTIONS

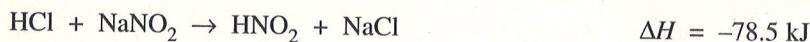
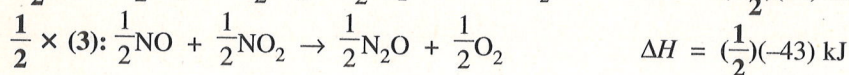
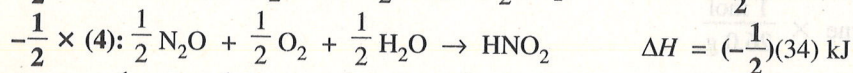
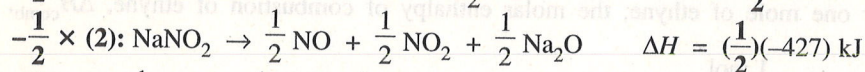
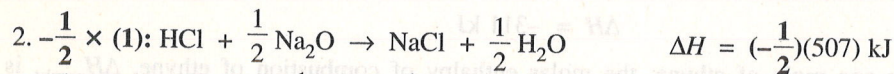
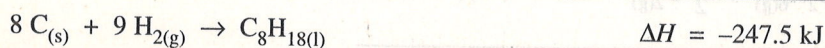
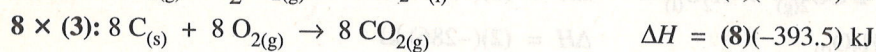
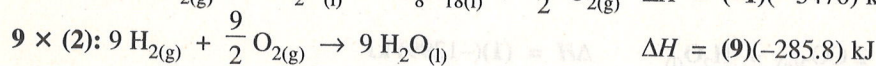
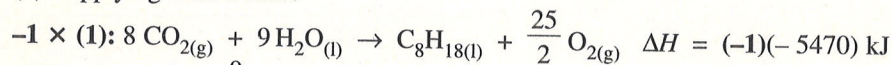
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Understanding Concepts

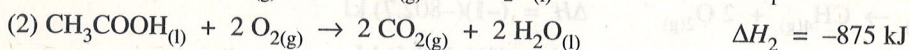
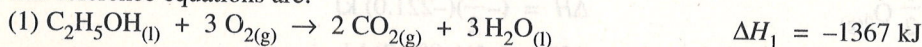
1. (a) The reference thermochemical equations are:



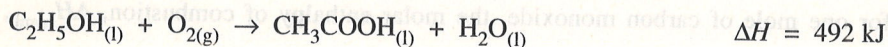
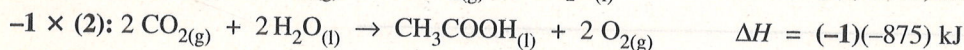
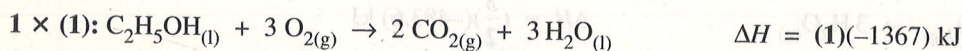
(b) Applying Hess's Law:



3. The reference equations are:

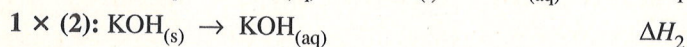
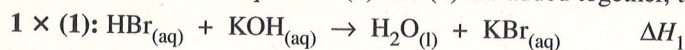


Applying Hess's Law:



Applying Inquiry Skills

4. (a) When reference equations (1) and (2) are added together, the result is reference equation (3).



(b) In all three experiments, assume 200.0 mL of solution is 200.0 g water.

Experiment 1:

$$q_{\text{water}} = mc\Delta T \\ = 200.0 \text{ g} \times 4.18 \text{ J/g}\cdot^{\circ}\text{C} \times (22.5 - 20.0)^{\circ}\text{C}$$

$$q_{\text{water}} = 2.09 \text{ kJ}$$

$$n\Delta H_1 = q_{\text{water}} \\ = 2.09 \text{ kJ}$$

$$\text{amount of KOH, } n = CV \\ = 1.00 \text{ mol/L} \times 0.100 \text{ L}$$

$$n = 0.100 \text{ mol}$$

$$\Delta H_1 = \frac{q}{n} \\ = \frac{2.09 \text{ kJ}}{0.100 \text{ mol}}$$

$$\Delta H_1 = 20.9 \text{ kJ/mol KOH}$$

Because the reaction is exothermic and is written for one mole of KOH, ΔH_1 is -20.9 kJ .

Experiment 2:

$$q_{\text{water}} = mc\Delta T \\ = 200.0 \text{ g} \times 4.18 \text{ J/g}\cdot^{\circ}\text{C} \times (24.1 - 20.0)^{\circ}\text{C}$$

$$q_{\text{water}} = 3.43 \text{ kJ}$$

$$n\Delta H_2 = q_{\text{water}} \\ = 3.43 \text{ kJ}$$

$$\text{molar mass of KOH, } M = 56.1 \text{ g/mol}$$

$$\text{amount of KOH, } n = 5.61 \text{ g} \times \frac{1 \text{ mol}}{56.1 \text{ g}}$$

$$n = 0.100 \text{ mol}$$

$$\Delta H_2 = \frac{q}{n} \\ = \frac{3.43 \text{ kJ}}{0.100 \text{ mol}}$$

$$\Delta H_2 = 34.3 \text{ kJ/mol KOH}$$

Because the reaction is exothermic and is written for one mole of KOH, ΔH_2 is -34.3 kJ .

Experiment 3:

$$q_{\text{water}} = mc\Delta T \\ = 200.0 \text{ g} \times 4.18 \text{ J/g}\cdot^{\circ}\text{C} \times (26.7 - 20.0)^{\circ}\text{C}$$

$$q_{\text{water}} = 5.60 \text{ kJ}$$

$$n\Delta H_{\text{reaction}} = q_{\text{water}} \\ = 5.60 \text{ kJ}$$

$$\text{amount of KOH, } n = 5.61 \text{ g} \times \frac{1 \text{ mol}}{56.1 \text{ g}}$$

$$n = 0.100 \text{ mol}$$

$$\Delta H_3 = \frac{q}{n} \\ = \frac{5.60 \text{ kJ}}{0.100 \text{ mol}}$$

$$\Delta H_3 = 56.0 \text{ kJ/mol KOH}$$

Because the reaction is exothermic and is written for one mole of KOH, ΔH_3 is -56.0 kJ/mol .

$$\begin{aligned}
 \text{(c) Experimental percentage error} &= \frac{|\Delta H_1 + \Delta H_2 - \Delta H_3|}{\Delta H_3} \times 100\% \\
 &= \frac{|(20.9 + 34.3) - 56.0|}{56.0} \times 100\% \\
 &= 1.4\%
 \end{aligned}$$

5.5 STANDARD ENTHALPIES OF FORMATION

PRACTICE

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Understanding Concepts

- $6 \text{C}_{(s)} + 3 \text{H}_{2(g)} \rightarrow \text{C}_6\text{H}_{6(l)}$
 - $\text{K}_{(s)} + \frac{1}{2} \text{Br}_{2(l)} + \frac{3}{2} \text{O}_{2(g)} \rightarrow \text{KBrO}_{3(s)}$
 - $6 \text{C}_{(s)} + 6 \text{H}_{2(g)} + 3 \text{O}_{2(g)} \rightarrow \text{C}_6\text{H}_{12}\text{O}_{6(s)}$
 - $\text{Mg}_{(s)} + \text{H}_{2(g)} + \text{O}_{2(g)} \rightarrow \text{Mg(OH)}_{2(s)}$

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Understanding Concepts

- $\text{C}_5\text{H}_{12(l)} + 8 \text{O}_{2(g)} \rightarrow 5 \text{CO}_{2(g)} + 6 \text{H}_2\text{O}_{(l)}$

$$\Delta H_{f(\text{C}_5\text{H}_{12(l)})}^\circ = -173.5 \text{ kJ/mol}$$

$$\Delta H_{f(\text{O}_{2(g)})}^\circ = 0 \text{ kJ/mol}$$

$$\Delta H_{f(\text{CO}_{2(g)})}^\circ = -393.5 \text{ kJ/mol}$$

$$\Delta H_{f(\text{H}_2\text{O}_{(l)})}^\circ = -285.8 \text{ kJ/mol}$$

$$\Delta H = 5 \Delta H_{f(\text{CO}_{2(g)})}^\circ + 6 \Delta H_{f(\text{H}_2\text{O}_{(l)})}^\circ - 1 \Delta H_{f(\text{C}_5\text{H}_{12(l)})}^\circ - 8 \Delta H_{f(\text{O}_{2(g)})}^\circ$$

$$= 5(-393.5) + 6(-285.8) - 1(-173.5) - 8(0)$$

$$\Delta H = -3509 \text{ kJ}$$
 - $\text{Fe}_2\text{O}_{3(s)} + 3 \text{CO}_{(g)} \rightarrow 2 \text{Fe}_{(s)} + 3 \text{CO}_{2(g)}$

$$\Delta H_{f(\text{Fe}_2\text{O}_{3(s)})}^\circ = -824.2 \text{ kJ/mol}$$

$$\Delta H_{f(\text{CO}_{(g)})}^\circ = -110.5 \text{ kJ/mol}$$

$$\Delta H_{f(\text{Fe}_{(s)})}^\circ = 0 \text{ kJ/mol}$$

$$\Delta H_{f(\text{CO}_{2(g)})}^\circ = -393.5 \text{ kJ/mol}$$

$$\Delta H = 2 \Delta H_{f(\text{Fe}_{(s)})}^\circ + 3 \Delta H_{f(\text{CO}_{2(g)})}^\circ - 1 \Delta H_{f(\text{Fe}_2\text{O}_{3(s)})}^\circ - 3 \Delta H_{f(\text{CO}_{(g)})}^\circ$$

$$= 2(0) + 3(-393.5) - 1(-824.2) - 3(-110.5)$$

$$\Delta H = -24.8 \text{ kJ}$$
- $\text{C}_6\text{H}_{12(l)} + 9 \text{O}_{2(g)} \rightarrow 6 \text{CO}_{2(g)} + 6 \text{H}_2\text{O}_{(l)}$

$$\Delta H_{\text{comb}(\text{C}_6\text{H}_{12(l)})}^\circ = -3824 \text{ kJ/mol}$$

$$\Delta H_{f(\text{O}_{2(g)})}^\circ = 0 \text{ kJ/mol}$$

$$\Delta H_{f(\text{CO}_{2(g)})}^\circ = -393.5 \text{ kJ/mol}$$

$$\Delta H_{f(\text{H}_2\text{O}_{(l)})}^\circ = -285.8 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{comb}(\text{C}_6\text{H}_{12(\text{l})})} = 6 \Delta H^\circ_{\text{f}(\text{CO}_{2(\text{g})})} + 6 \Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(\text{l})})} - 1 \Delta H^\circ_{\text{f}(\text{C}_6\text{H}_{12(\text{l})})} - 9 \Delta H^\circ_{\text{f}(\text{O}_{2(\text{g})})}$$

$$-3824 = 6(-393.5) + 6(-285.8) - 1(\Delta H^\circ_{\text{f}(\text{C}_6\text{H}_{12(\text{l})})}) - 9(0)$$

$$\Delta H^\circ_{\text{f}(\text{C}_6\text{H}_{12(\text{l})})} = 6(-393.5) + 6(-285.8) - 9(0) + 3824 \text{ kJ}$$

$$\Delta H^\circ_{\text{f}(\text{C}_6\text{H}_{12(\text{l})})} = -252 \text{ kJ/mol}$$

The standard enthalpy of combustion of liquid cyclohexane is -252 kJ/mol .

4. (a) $\Delta H^\circ_{\text{f}(\text{CH}_4(\text{g}))} = -74.4 \text{ kJ/mol}$

$$\Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(\text{l})})} = -285.8 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{CO}_{(\text{g})})} = -110.5 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{H}_2(\text{g}))} = 0 \text{ kJ/mol}$$

$$\Delta H = 1 \Delta H^\circ_{\text{f}(\text{CO}_{(\text{g})})} + 3 \Delta H^\circ_{\text{f}(\text{H}_2(\text{g}))} - 1 \Delta H^\circ_{\text{f}(\text{CH}_4(\text{g}))} + 1 \Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(\text{l})})}$$

$$= 1(-110.5) + 3(0) - (1(-74.4) + 1(-285.8))$$

$$\Delta H = 249.7 \text{ kJ } \mathbf{205.7}$$

ΔH for the reaction is 249.7 kJ .

(b) $\Delta H^\circ_{\text{f}(\text{CO}_{(\text{g})})} = -110.5 \text{ kJ/mol}$

$$\Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(\text{l})})} = -285.8 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{CO}_{2(\text{g})})} = -393.5 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{H}_2(\text{g}))} = 0 \text{ kJ/mol}$$

$$\Delta H = 1 \Delta H^\circ_{\text{f}(\text{CO}_{2(\text{g})})} + 1 \Delta H^\circ_{\text{f}(\text{H}_2(\text{g}))} - 1 \Delta H^\circ_{\text{f}(\text{CO}_{(\text{g})})} - 1 \Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(\text{l})})}$$

$$= 1(-393.5) + 1(0) - 1(-110.5) - 1(-285.8)$$

$$\Delta H = 2.8 \text{ kJ}$$

ΔH for the reaction is 2.8 kJ .

(c) $\Delta H^\circ_{\text{f}(\text{N}_2(\text{g}))} = 0 \text{ kJ/mol}$

$$\Delta H^\circ_{\text{f}(\text{H}_2(\text{g}))} = 0 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{NH}_3(\text{g}))} = -45.9 \text{ kJ/mol}$$

$$\Delta H = 2 \Delta H^\circ_{\text{f}(\text{NH}_3(\text{g}))} - 1 \Delta H^\circ_{\text{f}(\text{N}_2(\text{g}))} - 3 \Delta H^\circ_{\text{f}(\text{H}_2(\text{g}))}$$

$$= 2(-45.9) - 1(0) - 3(0)$$

$$\Delta H = -91.8 \text{ kJ}$$

ΔH for the reaction is -91.8 kJ .

5. (a) $\Delta H^\circ_{\text{f}(\text{NH}_3(\text{g}))} = -45.9 \text{ kJ/mol}$

$$\Delta H^\circ_{\text{f}(\text{O}_2(\text{g}))} = 0 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{NO}_{(\text{g})})} = 90.2 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(\text{l})})} = -285.8 \text{ kJ/mol}$$

$$\Delta H = 4 \Delta H^\circ_{\text{f}(\text{NO}_{(\text{g})})} + 6 \Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(\text{l})})} - 4 \Delta H^\circ_{\text{f}(\text{NH}_3(\text{g}))} - 5 \Delta H^\circ_{\text{f}(\text{O}_2(\text{g}))}$$

$$= 4(+90.2) + 6(-285.8) - 4(-45.9) - 5(0)$$

$$\Delta H = -1170.4 \text{ kJ}$$

ΔH for the reaction is -1170.4 kJ .

(b) $\Delta H^\circ_{\text{f}(\text{NO}_{(\text{g})})} = 90.2 \text{ kJ/mol}$

$$\Delta H^\circ_{\text{f}(\text{O}_2(\text{g}))} = 0 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{NO}_2(\text{g}))} = 33.2 \text{ kJ/mol}$$

$$\Delta H = 2 \Delta H_{f(\text{NO}_{2(\text{g})})}^{\circ} - 2 \Delta H_{f(\text{NO}_{(\text{g})})}^{\circ} - 1 \Delta H_{f(\text{O}_{2(\text{g})})}^{\circ}$$

$$= 2(+33.2) - 2(+90.2) - 1(0)$$

$$\Delta H = -114.2 \text{ kJ}$$

ΔH for the reaction is -114.2 kJ .

(c) $\Delta H_{f(\text{NO}_{2(\text{g})})}^{\circ} = 33.2 \text{ kJ/mol}$

$$\Delta H_{f(\text{H}_2\text{O}_{(\text{l})})}^{\circ} = -285.8 \text{ kJ/mol}$$

$$\Delta H_{f(\text{HNO}_{3(\text{l})})}^{\circ} = -174.1 \text{ kJ/mol}$$

$$\Delta H_{f(\text{NO}_{(\text{g})})}^{\circ} = 90.2 \text{ kJ/mol}$$

$$\Delta H = 2 \Delta H_{f(\text{HNO}_{3(\text{l})})}^{\circ} + 1 \Delta H_{f(\text{NO}_{(\text{g})})}^{\circ} - 3 \Delta H_{f(\text{NO}_{2(\text{g})})}^{\circ} - 1 \Delta H_{f(\text{H}_2\text{O}_{(\text{l})})}^{\circ}$$

$$= 2(-174.1) + 1(+90.2) - 3(+33.2) - 1(-285.8)$$

$$\Delta H = -71.8 \text{ kJ}$$

ΔH for the reaction is -71.8 kJ .

Making Connections

6. (a) raw materials + light energy + fertilizer (chemical energy) \rightarrow plant (chemical energy) \rightarrow prepared food (chemical energy) \rightarrow blood sugars (chemical energy) \rightarrow body tissue (chemical energy), movement (kinetic energy), or heat loss (thermal energy)
- (b) The advantage of using fertilizers to increase crop yield is that, although we input energy to make the fertilizers, the energy we gain back is now in a form that is usable by our bodies. The question of whether or not we should fertilize our crops goes far beyond a consideration of the amount of energy used to produce the fertilizer. It should include a weighing of the risks and benefits of both fertilizing (e.g., possibly polluting the ground water) and not fertilizing (e.g., producing less food).

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Making Connections

7. (a) $\Delta H_{f(\text{NH}_{3(\text{g})})}^{\circ} = -45.9 \text{ kJ/mol}$

$$\Delta H_{f(\text{HNO}_{3(\text{l})})}^{\circ} = -174.1 \text{ kJ/mol}$$

$$\Delta H_{f(\text{NH}_4\text{NO}_{3(\text{s})})}^{\circ} = -365.6 \text{ kJ/mol}$$

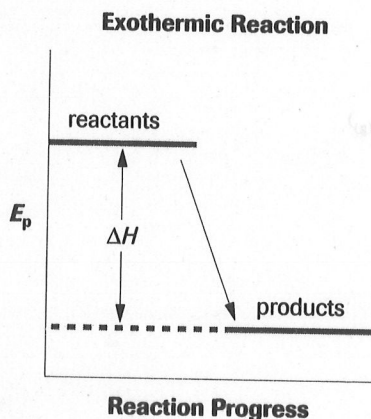
$$\Delta H = 1 \Delta H_{f(\text{NH}_4\text{NO}_{3(\text{s})})}^{\circ} - 1 \Delta H_{f(\text{NH}_{3(\text{g})})}^{\circ} - 1 \Delta H_{f(\text{HNO}_{3(\text{l})})}^{\circ}$$

$$= 1(-365.6) - 1(-45.9) - 1(-174.1)$$

$$\Delta H = -145.6 \text{ kJ}$$

ΔH for the reaction is -145.6 kJ .

(b)



(c) molar mass of NH_4NO_3 , $M = 80.0 \text{ g/mol}$

$$\text{amount of } \text{NH}_4\text{NO}_3, n = 5.0 \times 10^7 \text{ g} \times \frac{1 \text{ mol}}{80.0 \text{ g}}$$

$$n = 6.25 \times 10^5 \text{ mol}$$

$$q = \Delta H$$

$$= n\Delta H_r^\circ$$

$$= 6.25 \times 10^5 \text{ mol} \times 145.6 \text{ kJ}$$

$$q = 9.10 \times 10^7 \text{ kJ}$$

As the reaction is exothermic, $9.10 \times 10^7 \text{ kJ}$ of heat will be lost to the surroundings during the reaction.

8. $\Delta H_{\text{f}(\text{coal}_{(\text{s})})}^\circ = -396.4 \text{ kJ/mol}$

$$\Delta H_{\text{f}(\text{O}_{2(\text{g})})}^\circ = 0 \text{ kJ/mol}$$

$$\Delta H_{\text{f}(\text{CO}_{2(\text{g})})}^\circ = -393.5 \text{ kJ/mol}$$

$$\Delta H_{\text{f}(\text{H}_2\text{O}_{(\text{g})})}^\circ = -241.8 \text{ kJ/mol}$$

$$\begin{aligned} \Delta H &= 104 \Delta H_{\text{f}(\text{CO}_{2(\text{g})})}^\circ + 16 \Delta H_{\text{f}(\text{H}_2\text{O}_{(\text{g})})}^\circ - 2 \Delta H_{\text{f}(\text{coal}_{(\text{s})})}^\circ - 111 \Delta H_{\text{f}(\text{O}_{2(\text{g})})}^\circ \\ &= 104(-393.5) + 16(-241.8) - 2(-396.4) - 111(0) \end{aligned}$$

$$\Delta H = -4.40 \times 10^4 \text{ kJ or } -44.0 \text{ MJ}$$

$$\begin{aligned} \Delta H_{\text{comb}} &= \frac{\Delta H}{n} \\ &= \frac{-44.0 \text{ MJ}}{2 \text{ mol coal}} \end{aligned}$$

$$\Delta H = -22.0 \text{ MJ/mol coal}$$

molar mass of $\text{C}_{52}\text{H}_{16}\text{O}$, $M = 656 \text{ g/mol}$

$$\text{amount of } \text{C}_{52}\text{H}_{16}\text{O}, n = 1.00 \times 10^5 \text{ g} \times \frac{1 \text{ mol}}{656 \text{ g}}$$

$$n = 152 \text{ mol}$$

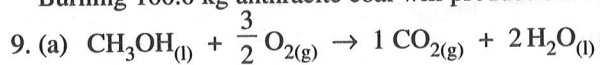
$$q = \Delta H$$

$$= n\Delta H_r^\circ$$

$$= 152 \text{ mol} \times 22.0 \text{ MJ/mol}$$

$$q = 3.34 \times 10^3 \text{ MJ}$$

Burning 100.0 kg anthracite coal will produce $3.34 \times 10^3 \text{ MJ}$ of thermal energy.



$$\Delta H_{\text{f}(\text{CH}_3\text{OH}_{(\text{l})})}^\circ = -239.1 \text{ kJ/mol}$$

$$\Delta H_{\text{f}(\text{O}_{2(\text{g})})}^\circ = 0 \text{ kJ/mol}$$

$$\Delta H_{\text{f}(\text{CO}_{2(\text{g})})}^\circ = -393.5 \text{ kJ/mol}$$

$$\Delta H_{\text{f}(\text{H}_2\text{O}_{(\text{l})})}^\circ = -285.8 \text{ kJ/mol}$$

$$\begin{aligned} \Delta H &= 1 \Delta H_{\text{f}(\text{CO}_{2(\text{g})})}^\circ + 2 \Delta H_{\text{f}(\text{H}_2\text{O}_{(\text{l})})}^\circ - 1 \Delta H_{\text{f}(\text{CH}_3\text{OH}_{(\text{l})})}^\circ - 1.5 \Delta H_{\text{f}(\text{O}_{2(\text{g})})}^\circ \\ &= 1(-393.5) + 2(-285.8) - 1(-239.1) - 1.5(0) \end{aligned}$$

$$\Delta H = -726 \text{ kJ}$$

molar mass of CH_3OH , $M = 32.0 \text{ g/mol}$

$$\text{amount of } \text{CH}_3\text{OH}, n = 1.00 \times 10^3 \text{ g} \times \frac{1 \text{ mol}}{32.0 \text{ g}}$$

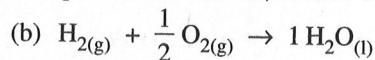
$$n = 31.3 \text{ mol}$$

$$q = \Delta H$$

$$= n\Delta H^\circ_{\text{comb}}$$

$$= 31.3 \text{ mol} \times 726 \text{ kJ/mol}$$

$$q = 2.27 \times 10^4 \text{ kJ, or } 22.7 \text{ MJ for one mole burned.}$$



$$\Delta H^\circ_{\text{f}(\text{H}_{2(\text{g})})} = 0 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{O}_{2(\text{g})})} = 0 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(\text{l})})} = -285.8 \text{ kJ/mol}$$

$$\Delta H = 1 \Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(\text{l})})} - 1 \Delta H^\circ_{\text{f}(\text{H}_{2(\text{g})})} - 0.5 \Delta H^\circ_{\text{f}(\text{O}_{2(\text{g})})}$$

$$= 1(-285.8) - 1(0) - 0.5(0)$$

$$\Delta H = -285.8 \text{ kJ}$$

$$\text{molar mass of H}_2, M = 2.02 \text{ g/mol}$$

$$\text{amount of H}_2, n = 1.00 \times 10^3 \text{ g} \times \frac{1 \text{ mol}}{2.02 \text{ g}}$$

$$n = 495 \text{ mol}$$

$$q = \Delta H$$

$$= n\Delta H^\circ_{\text{comb}}$$

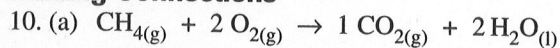
$$= 495 \text{ mol} \times 285.8 \text{ kJ/mol}$$

$$q = 1.42 \times 10^5 \text{ kJ, or } 142 \text{ MJ for one mole burned.}$$

(c) Both of these fuels produce less energy per mole than octane.

(d) Cost of the fuels and availability are two economic factors. Consumer safety concerns are also an issue since hydrogen is viewed as a more explosive gas.

Making Connections



$$\Delta H^\circ_{\text{f}(\text{CH}_{4(\text{g})})} = -74.4 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{O}_{2(\text{g})})} = 0 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{CO}_{2(\text{g})})} = -393.5 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(\text{l})})} = -285.8 \text{ kJ/mol}$$

$$\Delta H = 1 \Delta H^\circ_{\text{f}(\text{CO}_{2(\text{g})})} + 2 \Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(\text{l})})} - 1 \Delta H^\circ_{\text{f}(\text{CH}_{4(\text{g})})} - 2 \Delta H^\circ_{\text{f}(\text{O}_{2(\text{g})})}$$

$$= 1(-393.5) + 2(-285.8) - 1(-74.4) - 2(0)$$

$$\Delta H = -890.7 \text{ kJ/mol methane}$$

$$\text{Thus, } \Delta H^\circ_{\text{comb}} = -890.7 \text{ kJ/mol methane}$$

$$q_{\text{water}} = mc\Delta T$$

$$= 1.00 \times 10^5 \text{ g} \times 4.18 \text{ J/(g}\cdot^\circ\text{C)} \times (70 - 5)^\circ\text{C}$$

$$q_{\text{water}} = 2.7 \times 10^4 \text{ kJ}$$

$$n\Delta H^\circ_{\text{comb}} = q_{\text{water}}$$

$$n = \frac{q_{\text{water}}}{\Delta H^\circ_{\text{comb}}}$$

$$= \frac{2.7 \times 10^4 \text{ kJ}}{-890.7 \text{ kJ/mol}}$$

$$n = 30 \text{ mol}$$

$$m = n \times M$$

$$= 30 \text{ mol} \times 16.0 \text{ g/mol}$$

$$m = 480 \text{ g}$$

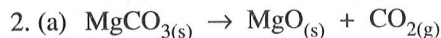
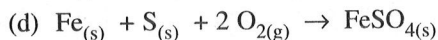
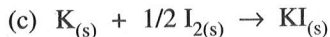
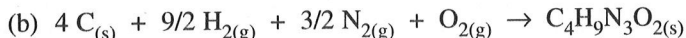
480 g of methane will heat 100 kg of water from 5°C to 70°C.

- (b) Insulating water pipes and setting the thermostat at a lower temperature are ways to conserve energy.
(c) Electricity and propane combustion are commonly used for heating water.

SECTION 5.5 QUESTIONS

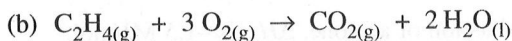
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Understanding Concepts



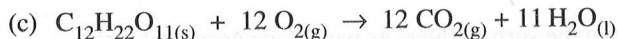
$$\begin{aligned}\Delta H &= 1 \Delta H_{f(\text{CO}_{2(g)})}^{\circ} + 1 \Delta H_{f(\text{MgO}_{(s)})}^{\circ} - 1 \Delta H_{f(\text{MgCO}_{3(s)})}^{\circ} \\ &= 1(-393.5) + 1(-601.6) - 1(-1095.8)\end{aligned}$$

$$\Delta H = 100.7 \text{ kJ}$$



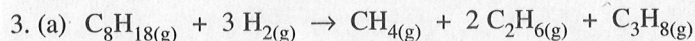
$$\begin{aligned}\Delta H &= 2 \Delta H_{f(\text{CO}_{2(g)})}^{\circ} + 2 \Delta H_{f(\text{H}_2\text{O}(l))}^{\circ} - 1 \Delta H_{f(\text{C}_2\text{H}_4(g))}^{\circ} - 3 \Delta H_{f(\text{O}_2(g))}^{\circ} \\ &= 2(-393.5) + 2(-285.8) - 1(+52.5) - 3(0)\end{aligned}$$

$$\Delta H = -1411 \text{ kJ}$$



$$\begin{aligned}\Delta H &= 12 \Delta H_{f(\text{CO}_{2(g)})}^{\circ} + 11 \Delta H_{f(\text{H}_2\text{O}(l))}^{\circ} - 1 \Delta H_{f(\text{sucrose}_{(s)})}^{\circ} - 12 \Delta H_{f(\text{O}_2(g))}^{\circ} \\ &= 12(-393.5) + 11(-285.8) - 1(-2225.5) - 12(0)\end{aligned}$$

$$\Delta H = -5640 \text{ kJ}$$

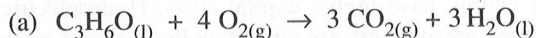


$$\begin{aligned}\Delta H &= 1 \Delta H_{f(\text{CH}_4(g))}^{\circ} + 2 \Delta H_{f(\text{C}_2\text{H}_6(g))}^{\circ} + 1 \Delta H_{f(\text{C}_3\text{H}_8(g))}^{\circ} - 1 \Delta H_{f(\text{C}_8\text{H}_{18(g)})}^{\circ} - 3 \Delta H_{f(\text{H}_2(g))}^{\circ} \\ &= 1(-74.4) + 2(-83.8) + 1(-104.7) - 1(-250.1) - 3(0)\end{aligned}$$

$$\Delta H = -96.6 \text{ kJ}$$

Applying Inquiry Skills

4. Prediction



$$\begin{aligned}\Delta H &= 3 \Delta H_{f(\text{CO}_{2(g)})}^{\circ} + 3 \Delta H_{f(\text{H}_2\text{O}(l))}^{\circ} - 1 \Delta H_{f(\text{acetone}_{(l)})}^{\circ} - 4 \Delta H_{f(\text{O}_2(g))}^{\circ} \\ &= 3(-393.5) + 3(-285.8) - 1(248.1) - 4(0)\end{aligned}$$

$$\Delta H = -1789.8 \text{ kJ or}$$

$$\Delta H_c = -1.79 \text{ MJ/mol acetone}$$

Analysis

$$\begin{aligned} \text{(b) } q_{\text{water}} &= mc\Delta T \\ &= 100.0 \text{ g} \times 4.18 \text{ J/g}\cdot^{\circ}\text{C} \times (25.0 - 20.0)^{\circ}\text{C} \\ q_{\text{water}} &= 2.09 \text{ kJ} \\ q_{\text{aluminum}} &= c\Delta T \\ &= 50.0 \text{ g} \times 0.91 \text{ J/g}\cdot^{\circ}\text{C} \times (25.0 - 20.0)^{\circ}\text{C} \\ q_{\text{aluminum}} &= 0.23 \text{ kJ} \\ q_{\text{total}} &= q_{\text{water}} + q_{\text{aluminum}} \\ q_{\text{water}} + q_{\text{aluminum}} &= 2.32 \text{ kJ} \\ n\Delta H_{\text{c}} &= q_{\text{total}} \\ m_{\text{acetone}} &= 0.092 \text{ g} \\ M_{\text{acetone}} &= 58.0 \text{ g} \\ n_{\text{acetone}} &= 0.092 \text{ g} \times \frac{1 \text{ mol}}{58.0 \text{ g}} \\ n_{\text{acetone}} &= 0.0016 \text{ mol} \\ \Delta H_{\text{c}} &= \frac{q_{\text{total}}}{n} \\ &= \frac{2.32 \text{ kJ}}{0.0016 \text{ mol}} \\ \Delta H_{\text{c}} &= 1.5 \text{ MJ/mol acetone} \end{aligned}$$

Because the reaction is exothermic, the molar enthalpy of combustion of acetone, ΔH_{c} , is -1.5 MJ/mol .

Evaluation

$$\begin{aligned} \text{(c) Percentage error} &= \frac{|(1.5 - 1.79)|}{1.79} \times 100\% \\ &= 16\% \end{aligned}$$

- (d) The percentage error suggests that heat has been lost to the surroundings as part of experimental error.
(e) If heat is lost to the surroundings, then the observed temperature change in the water and calorimeter, the calculated q values, and the experimental ΔH all will be smaller than predicted.

5.6 THE ENERGY DEBATE

PRACTICE

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Understanding Concepts

- (a) hydroelectric power, nuclear power, burning fossil fuels
(b) All produce power by driving turbines: nuclear and fossil fuels heat water to drive steam turbines, whereas hydroelectric power uses falling water. Hydroelectric and nuclear energy have higher capital costs. Hydroelectric is relatively environmentally benign, fossil fuel burning produces the greatest amount of pollution, and nuclear energy poses the greatest safety risk.

Making Connections

- (Sample answer) The CANDU system uses a heavy water moderator and ordinary uranium fuel, whereas some other systems use ordinary water as a moderator and enriched uranium fuel.